

# Chpt 7

P 215 Q 2 a, c, e, g  
3 a, c, e  
4 a - g

219 Q 2

228 Q 3

233 Q 1, 5

235 Q 10, 14

236 Q 26, 27

237 Q 33, 35, 37.

## Chpt 7 Test starts Here!

### Chpt 7: Language of Chemistry

$\text{CO}_2$  } formula itself means 1 molecule or 1 mole.

( $6.022 \times 10^{23}$  moles)

$\text{H}_2\text{O}$  }

↑ subscript → indicates # of constituent atoms

ex → 2 H-atoms & 1 O-atom

- 1 mole of  $\text{H}_2\text{O}$  ⇒ 2 moles of H-atoms

\* Once correct subscript are made they CAN NOT be changed! \*

[same w/ superscript]

### Making Ionic Compounds

#### Steps

ex.  $\text{Ca}^{2+}$   $\text{Cl}^-$

→ cations are always 1st

• step back... ionic compound →  $\text{NaCl}$

$\text{Na}^+ \rightarrow \text{Cl}^-$

→ net charge = 0 ~~★~~ always!!

ex. (cont)  $\text{CaCl}_2$  (need 2 chlorides to balance out 2+ charge of calcium)

ex.

$\text{Br}^-$

$\text{CO}_3^{2-}$

$\text{Mg}^{2+}$

$\text{MgBr}_2$

$\text{MgCO}_3$

$\text{Al}^{3+}$

$\text{AlBr}_3$

$\text{Al}_2(\text{CO}_3)_3$

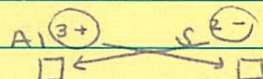
$\text{NH}_4^+$

$\text{NH}_4\text{Br}$

$(\text{NH}_4)_2\text{CO}_3$

ex  $\text{Al}^{3+} + \text{S}^{2-} \rightarrow \text{Al}_2\text{S}_3$

Cross-cross method :



→  $\text{Al}_2\text{S}_3$

• if  $\text{Mg}_2(\text{CO}_3)_2 \rightarrow$  not reduced!

→ always reduce subscripts



01-25-06	$\text{Br}^-$	$\text{SO}_4^{2-}$	$\text{ClO}_3^-$
$\text{K}^+$	$\text{KBr}$	$\text{K}_2\text{SO}_4$	$\text{KClO}_3$
$\text{Ba}^{2+}$	$\text{BaBr}_2$	$\text{BaSO}_4$	$\text{Ba}(\text{ClO}_3)_2$
$\text{NH}_4^+$	$\text{NH}_4\text{Br}$	$(\text{NH}_4)_2\text{SO}_4$	$\text{NH}_4\text{ClO}_3$

Naming ionic compounds (the stock system) ~~DO NOT~~ use prefixes for ionic compounds!!

1. Name the cation

2. Name the anion

ex.  $\text{NaCl}$  → sodium chloride

ex.  $\text{NaOH}$  → sodium hydroxide

ex.  $\text{KBr}$  → potassium bromide

ex.  $\text{BaBr}_2$  → barium bromide

ex.  $\text{NH}_4\text{Br}$  → ammonium bromide

ex.  $\text{Ba}(\text{ClO}_3)_2$  → barium chlorate

\* ate ion → higher charge.

\* ic ion → higher charge.

ex.  $\text{FeCl}_3$  → iron(III) chloride

stock system → use Roman numerals

ex.  $\text{FeCl}_2$  → iron(II) chloride

• old system (use latin names)

ex.  $\text{FeCl}_3$  → ferric chloride

$\text{FeCl}_2$  → ferrous chloride

poly  
valent  
ions

more than 1  
charge.



- Naming ionic compounds:
  - metal + nonmetal
  - metal + polyatomic

01-27-06

Naming Molecular Compounds → 2 nonmetals.

$\text{CO}$  → carbon monoxide

$\text{CO}_2$  → carbon dioxide

$\text{As}_2\text{O}_5$  → diarsenic pentoxide

• Rules of Naming

Step 1: Name the 1<sup>st</sup> element

(if more than one → give it a numeric prefix)

Step 2: Name the 2<sup>nd</sup> element, w/ numeric prefix & -ide suffix:

ex.  $\text{N}_2\text{O}_5$  → dinitrogen pentoxide

ex.  $\text{SiF}_4$  → Silicon tetrafluoride

ex.  $\text{XeF}_6$  → Xenon hexafluoride.

but beware...

Numeric Prefixes (drop vowels when they go like this)

• 1 → mono ( ) • 2 → di • 3 → tri • 4 → tetra • 5 → penta (a)

• 6 → hexa (a) • 7 → hepta • 8 → octa (a) • 9 → nona (a) • 10 → dec (a)

Naming Acids all start w/ Hydrogen.

- Types of acids → 1) Binary (H w/ another element)
- 2) Oxyacids (H w/ a polyatomic oxygen)

• Binary Acid:

ex  $\text{HCl}$  → hydrochloric acid.

→ hydro (prefix) + element name + -ic suffix = an acid.

ex  $\text{HF}$  → hydrofluoric acid

ex  $\text{HBr}$  → hydrobromic acid

• Oxyacids: ex  $\text{HNO}_3$  → Nitric acid

ex  $\text{H}_2\text{SO}_4$  → Sulfuric acid.

→ take anion name

- if it ends in -ate → change suffix to -ic } acid
- if it ends in -ite → change suffix to -ous }

ex  $\text{H}_2\text{CrO}_4$  → Chromic Acid

ex.  $\text{HNO}_2$  → Nitrous Acid



01-31-06

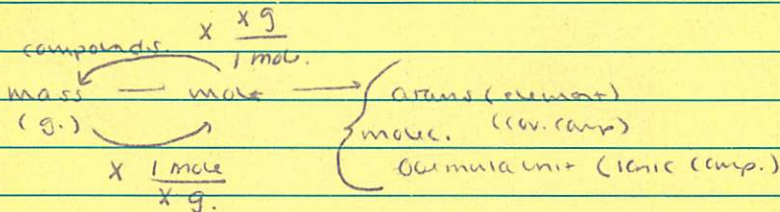
Mol concept 1 mol =  $6.022 \times 10^{23}$  "things"

• for elements ... 1 mol element =  $6.022 \times 10^{23}$  atoms

• for compounds ... Opt. a) 1 mole molecular (covalent bonds) =  $6.022 \times 10^{23}$  molecule (mole.)

Opt. b) 1 mole ionic (comp.) =  $6.022 \times 10^{23}$  formula units (fu)

• mol. concept of w/ compounds



for elements!  
ex. Mass → mole.

$$11.5 \text{ g Na} \times \frac{1 \text{ mol}}{23.0 \text{ g Na}} = 0.500 \text{ mol Na}$$

atomic mass expressed in g/mole  $\Rightarrow$  molar mass

for compounds

ex.  $6.7 \text{ g NaCl} = \underline{\hspace{2cm}} \text{ mol of NaCl}$

$$6.7 \text{ g NaCl} \times \frac{1 \text{ mol}}{\underline{\hspace{2cm}} \text{ g NaCl}}$$

g NaCl

← add their masses together!

• molar mass = the total mass <sup>of each atom</sup> present

(for comp.)



02-02-06

• 2.10 g  $\text{KMnO}_4$  → find: atoms of O

(39) + (55) + (4)(16) → 158

• 2.10 g  $\text{KMnO}_4$  •  $\frac{158 \text{ g KMnO}_4}{1 \text{ mol KMnO}_4} \cdot \frac{6.022 \times 10^{23} \text{ O atoms}}{1 \text{ mol KMnO}_4} \cdot \frac{8.00 \times 10^{21}}{8.00 \times 10^{21}} \text{ O atoms}$

• atoms of O...  $\frac{2.10 \text{ g KMnO}_4}{158 \text{ g KMnO}_4} \cdot \frac{1 \text{ mol KMnO}_4}{1 \text{ mol KMnO}_4} \cdot \frac{4 \text{ mol O}}{1 \text{ mol KMnO}_4} \cdot \frac{6.022 \times 10^{23} \text{ atoms O}}{1 \text{ mol O}}$

ex.  $\text{H}_2\text{O}$

2 atoms 1 atom

$1.2 \times 10^{24} \text{ atoms H} \cdot \frac{6.022 \times 10^{23} \text{ atoms O}}{1 \text{ mol O}}$

2 mol 1 mol

$\frac{1.2 \times 10^{24} \text{ atoms H}}{2 \text{ mol H}} \cdot \frac{1 \text{ mol H}_2\text{O}}{2 \text{ mol H}} \cdot \frac{1 \text{ mol O}}{1 \text{ mol H}_2\text{O}} \cdot \frac{6.022 \times 10^{23} \text{ atoms O}}{1 \text{ mol O}} = 3.20 \times 10^{23} \text{ atoms O}$

• atoms # of p. from O...

- each atom of O → 8 protons...

$3.20 \times 10^{23} \text{ atoms O} \cdot \frac{8 \text{ p.p.}}{1 \text{ atom O}} = 2.56 \times 10^{24} \text{ p.p.}$



02-03-06

Percent composition:

total mass of element in compound / total mass of the compound = 100%

ex.  $H_2O$

assumption

$\% H = 11.1\%$

$\% O = 88.9\%$

wrong!

$\% = \frac{\text{mass element}}{\text{1-Comp}} \times 100$

by mass!

$$\text{ex } H_2O \rightarrow \% O = \frac{16 \frac{g}{mol}}{18 \frac{g}{mol}} \cdot 100\% = 88.9\% O$$

$$\% H = 100\% - 88.9\% = 11.1\% H$$

How many grams of O are in 7.20 g of sulfuric acid?

$(H_2SO_4) \Rightarrow$  molar mass = 98  $\frac{g}{mol}$

molar mass O  $\rightarrow 64 \frac{g}{mol}$

7.20 g  $H_2SO_4$

$$\% O = \frac{64 \frac{g}{mol}}{98 \frac{g}{mol}} \cdot 100\% = 65.3\% O$$

$$65.3\% O = 0.653$$

$$0.653 \frac{g O}{g H_2SO_4} \cdot 7.20 g H_2SO_4 = 4.70 g O$$

- harder way to solve problem  $\rightarrow$  go through the moles.

Lab on composition of hydrates.

ex.  $CuSO_4 \cdot xH_2O$  [blue]

(heat)



$CuSO_4 + H_2O \uparrow$  [white]

water is bonded to compound  
 $\Rightarrow$  makes it more stable.

hydrate compound

anhydrous compound.

ex  $BF_3 \rightarrow$  readily reacts w/ water.

- compare two masses  $\rightarrow$  to find value of "x" in formula



02-08-06

## Empirical formulas & Molecular formulas

[most reduced]

• empirical formula → the simplest (smallest whole # ratios) formula possible

ex  $\text{NaCl}$ ,  $\text{MgSO}_4$ ,  $\text{CaCl}_2$ , or  $\text{CH}_2\text{O}$

molecular compound.

are ionic compounds

- all ionic compounds have empirical formulas.

• molecular formula - whole # multiple of the empirical formula based on molar mass.

ex  $\text{C}_6\text{H}_{12}\text{O}_6$  ← extremely different!

[its empirical formula is  $\text{CH}_2\text{O}$ ]

- molar mass of  $\text{CH}_2\text{O}$  =  $30 \frac{\text{g}}{\text{mole}}$

- molar mass of  $\text{C}_6\text{H}_{12}\text{O}_6$  =  $(6)30 = 180 \frac{\text{g}}{\text{mole}}$

EX. 1: 32.3% Na      22.65% S      44.99% O. → find empirical formula.

\* law of definite composition → same ratios of % composition no matter the source or size of a sample.

→ can use any #, but picking 100 only makes it easier.

Step 1: assume 100 g. of compound → what's the mass of each element present?

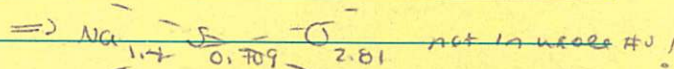
32.3 g Na, 22.65 g S, 44.99 g O.

Step 2: convert to moles.

$$\bullet \frac{32.3 \text{ g Na}}{23 \frac{\text{g Na}}{\text{mol Na}}} = 1.41 \text{ mol Na}$$

$$\bullet \frac{22.65 \text{ g S}}{32 \frac{\text{g S}}{\text{mol S}}} = 0.709 \text{ mol S}$$

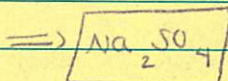
$$\bullet \frac{44.99 \text{ g O}}{16 \frac{\text{g O}}{\text{mol O}}} = 2.81 \text{ mol O.}$$



Step 3: divide by the lowest # of moles present.

$$\bullet \frac{1.41 \text{ mol Na}}{0.709} = 1.99 \frac{\text{mol Na}}{\text{mol S}} \Rightarrow 2 \quad \bullet \frac{0.709 \text{ mol S}}{0.709} = 1 \text{ mol S}$$

$$\bullet \frac{2.81 \text{ mol O}}{0.709} = 3.96 \text{ mol O.} \Rightarrow 4$$





02-08-06 Empirical formulas & Molecular formulas (cont)

Ex. 2: compound  $\rightarrow$  34.0 g/mol ; 0.44 g H & 6.92 g O  $\rightarrow$  what's the molecular formula?

Step 1  $\rightarrow$  already given the grams.

Step 2: convert to moles.

$$\frac{0.44 \text{ g H}}{1 \text{ g H}} \\ \text{mol H}$$

$$\frac{6.92 \text{ g O}}{16 \text{ g O}} \\ \text{mol O}$$

$$= \frac{0.44 \text{ mol H}}{0.44}$$

$$= \frac{0.43 \text{ mol O}}{0.44}$$

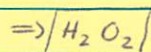
$$= 1 \text{ mol H}$$

$$= 1 \text{ mol O}$$

$\Rightarrow$  empirical formula  $\rightarrow$  HO

$\rightarrow$  find molecular - given molar mass = 34.0  $\frac{\text{g}}{\text{mol}}$

$$= \text{molar mass of } \rightarrow 17 \frac{\text{g}}{\text{mol}} \Rightarrow \frac{34.0 \frac{\text{g}}{\text{mol}}}{17.0 \frac{\text{g}}{\text{mol}}} = 2$$



• trick for step 3  $\rightarrow$  is 1.5 for moles  $\rightarrow$  double everything!